Chemistry 12

Acid-Base Equilibrium II

I 6-10

pg. 121-133

Name:
Date:
Block:

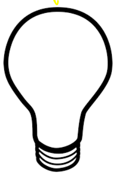
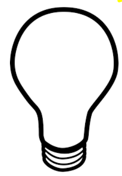
1. **Strengths of Acids and Bases**
2. **Relative Strengths of Brønsted-Lowry Acids and Bases**
3. **K_a, K_b**
4. **Ionization of Water**

Strengths of Acids and Bases

Demo: Consider two acid solutions – hydrochloric acid, HCl and acetic acid, CH_3COOH.

strong acid

weak acid

HCl $\rightarrow H^+ + Cl^-$ 1.0M 1.0M 1.0M	**CH_3COOH** $\rightleftharpoons H^+ + CH_3COO^-$ 1.0M <1.0M <1.0M
	

- more ions
- better ability to conduct electricity

strong

$[SA] = [H_3O^+]$

ex. $[HCl] = [H_3O^+]$

In order for a solution to conduct electricity, ions must be present.

A strong acid (or base) is a substance that completely ionizes in aqueous solution. $\rightarrow$

A weak acid (or bases) is a substance that partially ionizes in aqueous solution. $\rightleftharpoons$

HCl is a **strong** acid.

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

pg. 6 of data booklet

not listed

- The reaction goes to completion – therefore, $\rightarrow$ is used.
- Other strong acids: $HClO_4$, HI, HBr, HNO_3 and H_2SO_4.
- Strong bases include: NaOH, KOH, $Ca(OH)_2$, and $Mg(OH)_2$

CH_3COOH is a **weak** acid.

- The reaction does not go to completion – therefore $\leftrightharpoons$ is used.
- There are many weak acids and bases.

spectator ion = ions formed from strong acids

Classify the following as a strong acid (SA), weak acid (WA), strong base (SB), weak base (WB) or a spectator ion (S).

$HClO_4$ SA	HCOOH WA	NH_3 WB	CN^- WB
$Mg(OH)_2$ SB	HNO_2 WA	NO_3^- S	HNO_3 SA
HPO_4^{2-} WA / WB	F^- WB	Br^- S	Cl^- S
NH_4^+ WA	SO_3^{2-} WB	HSO_4^- WA	LiOH SB

acids bases

Relative Strengths of Brønsted-Lowry Acids and Bases

The higher the K_a or K_b value, the higher $[H_3O^+]$ or $[OH^-]$ respectively.

Acids are listed on the left of the table and their conjugate bases are listed on the right.

Example.

A student tests the electrical conductivity of 0.5 M solutions of the following: carbonic acid, methanoic acid, phenol acid and boric acid. Rank these solutions in order from most conductive to least conductive.

Most conductive — Least conductive

Methanoic > carbonic > boric > phenol

- stronger acid
- higher Ka value

When a BL-acid and base react, the position of the equilibrium is determined by their relative strengths.

Example:

$$NH_{3\,(aq)} + HCOOH_{(aq)} \leftrightharpoons NH_4^+{}_{(aq)} + HCOO^-{}_{(aq)}$$

B (stronger) A (stronger) A (weaker) B (weaker)

- equilibrium favours products
- $K_{eq} = \frac{P}{R}$

$K_{eq} > 1$

- Label the respective acids and bases in the above equation.
- According to the BL-acid/base strength table:
  - Determine which of the two acids is weaker or stronger.
  - Determine which of the two bases is weaker or stronger.
- **The equilibrium favours the direction in which the stronger acid and base react to form the weaker conjugate acid and base.**

stronger one will be the acid

Practice:

1. Predict whether reactants or products will be favoured when HCN reacts with HCO_3^-.

- What is the reaction? Identify the weaker/stronger acid, weaker/stronger base.

$$\underset{\underline{A}}{HCN} + \underset{B}{HCO_3^-} \rightleftharpoons \underset{\substack{B \\ \text{stronger}}}{CN^-} + \underset{\substack{\underline{A} \\ \text{stronger}}}{H_2CO_3}$$

$\boxed{\therefore \text{Favours reactants}}$

2. Predict whether reactants or products will be favoured when HSO_4^- reacts with $HC_2O_4^-$.

$$\underset{\underline{A}}{HSO_4^-} + \underset{B}{HC_2O_4^-} \rightleftharpoons \underset{\substack{B \\ \text{stronger}}}{SO_4^{2-}} + \underset{\substack{\underline{A} \\ \text{stronger}}}{H_2C_2O_4}$$

$\boxed{\therefore \text{Favours reactants}}$

a. Will the K_{eq} value be greater or less than 1?

$K_{eq} = \frac{P}{R}$ $\quad\quad K_{eq} < 1$

3. If 0.10 M HSO_3^- is mixed with 0.10 M $HC_2O_4^-$, which species will ***donate*** a proton?

A / B $\quad\quad$ A / B

compare strengths

acid

$\boxed{HC_2O_4^-}$

4. If 0.10 M $\cancel{HSO_4^-}$ is mixed with 0.10 M $HC_6H_5O_7^{2-}$, which species will ***accept*** a proton?

$\cancel{A}$ $\quad\quad$ A / B

Cannot be base

base

$\boxed{HC_6H_5O_7^{2-}}$

<u>**Strengths of Acids & Bases Worksheet**</u>
<u>**Hebden Workbook Pg. 125 # 21-24, Pg 133 # 38 - 46**</u>

K_a, K_b

acid

$$CH_3COOH_{(aq)} + \cancel{H_2O}_{(l)} \leftrightharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

In the acetic acid example above, because the reaction reaches an equilibrium, a K_{eq} expression can be written. This specific expression is called **the K_a expression** and K_a is called the **acid ionization constant**.

$$K_{eq} = K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.8 \times 10^{-5}$$

Similarly for weak bases,

base

$$NH_{3\,(aq)} + \cancel{H_2O}_{(l)} \leftrightharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

The K_b expression is:

$$K_{eq} = K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \text{need to calculate}$$

K_b is called the base ionization constant.

Practice:

1. HF is a weak acid.
   a. Write an equation showing how HF acts in solution.

   $$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + F^-_{(aq)}$$

   b. Write the K_a expression for HF.

   $$K_a = \frac{[H_3O^+][F^-]}{[HF]} = 3.5 \times 10^{-4}$$

   data booklet

2. The hydrogen oxalate ion is amphiprotic.
   a. Write an equation showing how this ion acts as an acid in solution. Write a K_a expression.

   $$HC_2O_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons C_2O_4^{2-}{}_{(aq)} + H_3O^+_{(aq)}$$

   $$K_a = \frac{[C_2O_4^{2-}][H_3O^+]}{[HC_2O_4^-]}$$

   b. Write an equation showing how this ion acts as a base in solution. Write a K_b expression.

   $$HC_2O_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_2C_2O_{4\,(aq)} + OH^-_{(aq)}$$

   $$K_b = \frac{[H_2C_2O_4][OH^-]}{[HC_2O_4^-]}$$

<u>Lab Simulation: Strengths of Acids and Bases</u>

Ionization of Water

Water is amphiprotic so it can donate and accept a proton ion.

- One water can donate a proton to another water molecule – this is called **autoionization.**

$$\underset{B}{H_2O}_{(l)} + \underset{A}{H_2O}_{(l)} \leftrightharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

What would be the equilibrium constant expression?

$$K_{eq} = K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

The equilibrium constant expression for the autoionization of water is called K_w, or the water ionization constant or ion product constant.

At equilibrium, $[H_3O^+] = [OH^-]$

(neutral)

$$H_2O_{(l)} + H_2O_{(l)} \leftrightharpoons \underset{x}{H_3O^+_{(aq)}} + \underset{x}{OH^-_{(aq)}}$$

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$(x)(x) = 1.0 \times 10^{-14}$$

$$x = 1.0 \times 10^{-7} = [H_3O^+] = [OH^-]$$

Both H_3O^+ and OH^- exist in all aqueous solution.

- If $[H_3O^+] < [OH^-]$, the solution is basic.
- If $[H_3O^+] = [OH^-]$, the solution is neutral.
- If $[H_3O^+] > [OH^-]$, the solution is acidic.

$[SA] = [H_3O^+]$

Example:

What is the $[H_3O^+]$ and $[OH^-]$ in 0.50M HCl? Is this solution acidic or basic?

- Is HCl a strong or weak acid?

- strong

$$\underset{0.50M}{HCl} + H_2O \xrightarrow{\quad} \boxed{\underset{0.50M}{H_3O^+}} + \cancel{Cl^-} \text{ spectator}$$

1:1

- Substitute known values into the K_w expression and solve.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{0.50}$$

$$= \boxed{2.0 \times 10^{-14} M}$$

$\therefore [H_3O^+] > [OH^-]$

solution is $\boxed{\text{acidic}}$

Practice 1:

Ksp!

strong base

Calculate the $[H_3O^+]$ and $[OH^-]$ in a **saturated** solution of magnesium hydroxide. Is this solution acidic or basic?

$Mg(OH)_2 \rightarrow \overset{s}{Mg^{2+}} + \overset{2s}{2OH^-}$

$K_{sp} = 5.6 \times 10^{-12} = 4s^3$

$s = 1.1 \times 10^{-4}$

$2s = [OH^-] = \boxed{2.2 \times 10^{-4} M}$

$K_w = [H_3O^+][OH^-]$

$[H_3O^+] = \frac{K_w}{[OH^-]}$

$= \frac{1.0 \times 10^{-14}}{2.2 \times 10^{-4}}$

$= \boxed{4.5 \times 10^{-11} M}$

$[OH^-] > [H_3O^+]$

solution is $\boxed{\text{basic}}$

Practice 2:

A student dissolved 1.42g of NaOH in 250. mL of solution. Calculate the resulting $[H_3O^+]$ and $[OH^-]$. Is this solution acidic or basic?

$[NaOH] = \frac{1.42 g}{0.250 L} \times \frac{1 mol}{40.0 g} = 0.142 M$

$NaOH \rightarrow Na^+ + \boxed{\begin{matrix} OH^- \\ 0.142M \end{matrix}}$

$0.142M$

$[H_3O^+] = \frac{K_w}{[OH^-]}$

$= \frac{1.0 \times 10^{-14}}{0.142}$

$= \boxed{7.0 \times 10^{-14} M}$

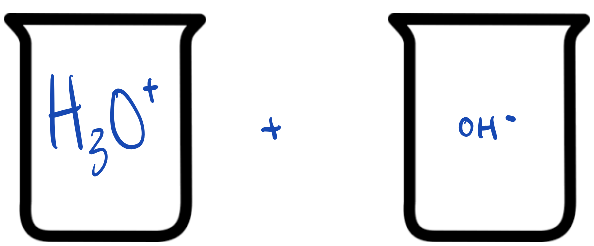
$[OH^-] > [H_3O^+]$

$\therefore$ solution is $\boxed{\text{basic}}$

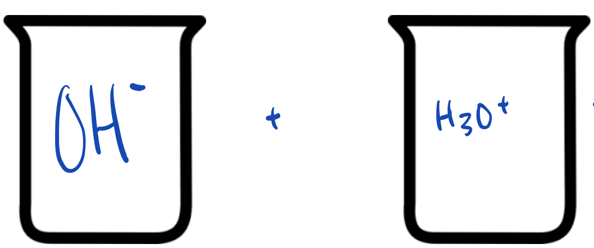
<u>$[H_3O^+]$ and $[OH^-]$ when 2 solutions are MIXED:</u>

When two solutions are mixed:

- The solutions <u>dilute</u> each other. ($c_1v_1 = c_2v_2$)
- The acid and base <u>neutralize</u> each other. new!



$= [H_3O^+]_i - [OH^-]_i = \boxed{[H_3O^+]_f \text{ left over}}$



$= [OH^-]_i - [H_3O^+]_i = \boxed{[OH^-]_f \text{ left over}}$

Practice 1: doesn't necessarily mean acidic solution!

What is the final $[H_3O^+]$ in a solution formed when 25 mL of 0.30 M HCl is added to 35 mL of 0.50 M NaOH?

Step 1 Dilution

- When two solutions are combined, both are diluted. Calculate the new concentrations of HCl and NaOH in the mixed solution.

$[SA] = [H_3O^+]$

$[HCl]$

$$C_1V_1 = C_2V_2$$

$$C_2 = \frac{(0.30)(25)}{(60)} = 0.125M = [H_3O^+]$$

$[NaOH]$

$$C_1V_1 = C_2V_2$$

$$C_2 = \frac{(0.50)(35)}{(60)} = 0.292M = [OH^-]$$

Step 2 Neutralization

- The hydronium ions and hydroxide ions will neutralize each other.
- Since there is more $\underline{OH^-}$, there will be $\underline{OH^-}$ left over. Calculate how much will be left over.

$$[OH^-]_f = [OH^-]_i - [H_3O^+]_i$$
$$= 0.292M - 0.125M$$
$$= 0.167M = \boxed{0.17M}$$

- Use K_w to calculate the hydronium ion concentrations from the hydroxide ion concentration.

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0\times10^{-14}}{0.167M} = \boxed{6.0\times10^{-14}M}$$

Practice 2:

Calculate the final $[H_3O^+]$ and $[OH^-]$ in a solution formed when 150. mL of 1.5 M HNO_3 is added to 250. mL of 0.80 M KOH.

$[HNO_3]$

$$C_2 = \frac{(1.5)(150)}{400}$$
$$= 0.5625M = [H_3O^+]$$

$[KOH]$

$$C_2 = \frac{(0.80)(250)}{400}$$
$$= 0.5000M = [OH^-]$$

$$[H_3O^+]_f = 0.5625 - 0.5000$$
$$= 0.0625M = \boxed{0.063M}$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{0.0625} = \boxed{1.6\times10^{-13}M}$$

Both acids!

$$[H_3O^+]_f = [H_3O^+]_i + [H_3O^+]_i$$

add!

Practice 3:

Calculate the resulting $[H_3O^+]$ and $[OH^-]$ when 18.4 mL of 0.105 M HBr is added to 22.3 mL of 0.256 M HCl.

[HBr]

$$C_2 = \frac{(0.105)(18.4)}{40.7}$$

$$= 0.047469\,M = [H_3O^+]$$

[HCl]

$$C_2 = \frac{(0.256)(22.3)}{40.7}$$

$$= 0.140265\,M = [H_3O^+]$$

$$[H_3O^+]_f = 0.047469 + 0.140265$$

$$= 0.01877 = \boxed{0.188\,M}$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{0.1877} = \boxed{5.3\times10^{-14}\,M}$$

Practice 4:

What mass of NaOH must be added to 500.0 mL of a solution of 0.020 M HI to obtain a solution with a final hydronium ion concentration of 0.0032 M?

$[H_3O^+]_i$

$[H_3O^+]_f$

$$[H_3O^+]_f = [H_3O^+]_i - [OH^-]_i$$

$$0.0032 = 0.020 - [OH^-]_i$$

$$[OH^-]_i = 0.0168\,M$$

$$0.5000\,L \times \frac{0.0168\,mol}{1\,L} \times \frac{40.0\,g}{1\,mol} = 0.336 = \boxed{0.34\,g}$$